

1. Refer to the phase diagram for carbon to answer the following questions.

a. What phase of carbon is present under ambient (1 atm, 298 K) conditions? **Graphite**

b. If the temperature is increased at a constant 1 atm, what phase change occurs first? **Graphite to vapor**

c. How many triple points exist in this diagram? **Two**

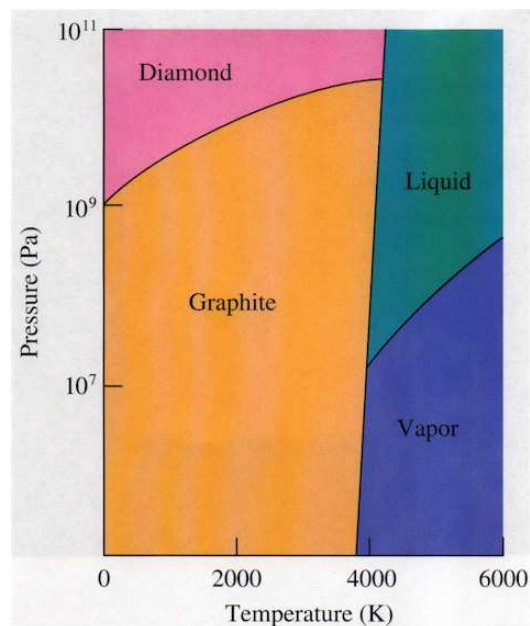
d. Which phases coexist at each triple point?

G, D, L and G, L, V

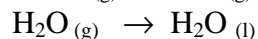
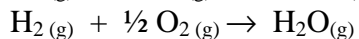
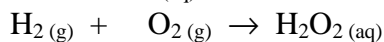
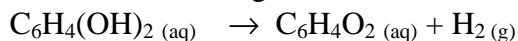
e. Assuming that density increases with pressure, which phase, graphite or diamond, is more dense? **Diamond**

f. Abrasive-grade diamond grit can be synthesized in the lab under high temperatures and pressures. At 3500 K, what is the minimum pressure needed to accomplish this feat?

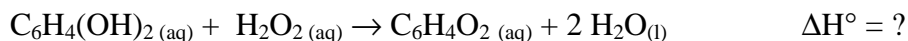
About 10^{10} Pa



2. Given the following reactions:



Calculate ΔH° for the conversion of hydroquinone to quinone, used by the bombardier beetle as an explosive defense mechanism:



By use of Hess' Law, reverse reaction 2, double reaction 3 and reaction 4, sum together to yield the overall desired equation. Accordingly, you must change the sign of ΔH for reaction 2 and multiply the ΔH values of reactions 3 and 4 by two. Then sum the ΔH values for the individual reactions. This is equal to ΔH for the overall reaction.

$$177.4 + 191.2 + 2(-241.8) + 2(-43.8) = -202.6 \text{ kJ/mol}$$

3. What is the difference in energy for 1.00 mol of an ideal gas heated from 300 K to 350 K in a 10.0 L flask? What is the molar heat capacity of the gas? Remember that $E_k = \frac{3}{2}RT$.

$$E_k = (2/3)(8.314 \text{ J/mol}\cdot\text{K})(50 \text{ K}) = 624 \text{ J}$$

$$624 \text{ J} / [(1 \text{ mol})\cdot(50 \text{ K})] = 12.47 \text{ J/mol}\cdot\text{K} \quad (= 3/2 R)$$

4. Consider 2.00 mol of an ideal gas held at a constant temperature 298 K. The sample is allowed to expand from 9.00 L to 36.00 L against a constant pressure of 1.00 atm.

a. What is the work done by the gas (remember 1 L·atm = 101 J)?

$$\text{work} = -P \cdot \Delta V = -(1 \text{ atm})(27 \text{ L})(101 \text{ J/L}\cdot\text{atm}) = -2.7 \text{ kJ}$$

b. What is ΔE for the gas?

$$\text{Because the temperature stayed constant } (\Delta T = 0) \text{ and } E_k = 3/2RT, \Delta E = 0$$

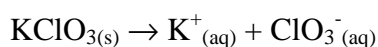
c. What is ΔH for the gas?

By definition, ΔH is the heat (q) at constant pressure; so $\Delta H = q$.

$$\Delta E = q + w = \Delta H + w$$

$$\Delta H = \Delta E - w = 0 - (-2.7 \text{ kJ}) = 2.7 \text{ kJ of heat absorbed}$$

5. When 1.00 g of potassium chlorate (KClO_3 ; MW = 122.6 g/mol) is dissolved in 50.0 g of water, the temperature drops from 25.00°C to 23.36°C (water has a specific heat capacity of 4.184 J/g°C). What is the change in enthalpy for the following reaction (which describes the dissolution of KClO_3 .)



$$q_{\text{lost}} = q_{\text{gained}}$$

$$q_{\text{lost}} = (50 \text{ g})(4.184 \text{ J/g}\cdot\text{°C})(1.64 \text{ °C}) = 343 \text{ J}$$

$$\Delta H = q = 343 \text{ J (but how much energy on a per mole basis?)}$$

$$(1 \text{ g})/(122.6 \text{ g/mol}) = 0.0082 \text{ mol}$$

$$\text{So } \Delta H = 343 \text{ J} / 0.0082 \text{ mol} = 42 \text{ kJ/mol}$$

6. A quantity of heat, 5 kJ, is added to a calorimeter that contains 100 g of water, causing the temperature of the water and calorimeter to rise from 23.5 °C to 34.4 °C. (a) What is the heat capacity of the calorimeter? (b) In a separate experiment, 30.0 g of an unknown metal initially at 80.0 °C is added to 100 g of water (in the same calorimeter), causing the temperature to rise from 20.0 °C to 20.9 °C. What is the identity of the metal?

specific heats (J/g·°C)

water: 4.184

aluminum: 0.901

copper: 0.385

silver: 0.235

gold: 0.129

a) $q_{\text{lost}} = q_{\text{H}_2\text{O}} + q_{\text{cal}}$

$$5000 \text{ J} = (100 \text{ g})(4.184 \text{ J/g}\cdot\text{°C})(10.9 \text{ °C}) + C_{\text{cal}} (10.9 \text{ °C})$$

$$C_{\text{cal}} = \text{heat capacity of the calorimeter} = 40.3 \text{ J/°C}$$

b) $q_{\text{lost}} = q_{\text{cold H}_2\text{O}} + q_{\text{cal}}$

$$m_{\text{met}} \cdot c_{\text{met}} \cdot \Delta T_{\text{met}} = m_{\text{H}_2\text{O}} \cdot c_{\text{H}_2\text{O}} \cdot \Delta T_{\text{H}_2\text{O}} + m_{\text{cal}} \cdot c_{\text{cal}} \cdot \Delta T_{\text{cal}}$$

$$m_{\text{met}} \cdot c_{\text{met}} \cdot \Delta T_{\text{met}} = m_{\text{H}_2\text{O}} \cdot c_{\text{H}_2\text{O}} \cdot \Delta T_{\text{H}_2\text{O}} + C_{\text{cal}} \cdot \Delta T_{\text{cal}}$$

$$(30 \text{ g})(c_{\text{met}})(59.1 \text{ °C}) = (100 \text{ g})(4.184 \text{ J/g}\cdot\text{°C})(0.9 \text{ °C}) + (40.3 \text{ J/°C})(0.9 \text{ °C})$$

$$(1773 \text{ g}\cdot\text{°C})(c_{\text{met}}) = 413 \text{ J}$$

$$c_{\text{met}} = 0.232 \text{ J/g}\cdot\text{°C} \text{ and the metal is silver}$$